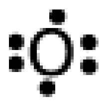
Set 19.0: Lewis Structures Part I

Skill 19.01: Draw Lewis structures for atoms
Skill 19.02: Explain the octet rule
Skill 19.03: Explain the formation single, double, and triple covalent bonds
Skill 19.04: Draw Lewis structures for atoms that follow the octet rule

Skill 19.01: Draw Lewis structures for atoms

Skill 19.01 Concepts

The electrons involved in chemical bonding are the valence electrons, those residing in the incomplete outer shell of an atom. The American chemist G.N. Lewis suggested a simple way of showing these valence electrons, which we now call Lewis electron-dot symbols or simply Lewis symbols. The Lewis symbol for an element consists of the chemical abbreviation for the element plus a dot for each valence electron. For example oxygen has the electron configuration $[He]2s^2 2p^4$; its Lewis symbol therefore shows six valence electrons. The dots are placed on the four sides of the atomic abbreviation.



Each side can accommodate up to two electrons. All four sides are equivalent; the placement of two electrons versus one is arbitrary. The number of valence electrons of any representative element is the same as the group number of the element in the periodic table.

Skill 19.01 Problem 1

Draw Lewis structures for the following atoms:				
(a) Ar	(b) Cl	(c) S	(d) P	(e) Si

Skill 19.02: Explain the octet rule

Skill 19.02 Concepts

Atoms often gain, lose, or share electrons so as to achieve the same number of electrons as the noble gas closest to them in the periodic table. Because all noble gases (except He) have eight valence electrons, many atoms undergoing reactions also end up with eight valence electrons. This observation has led the octet rule: atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons. An octet of electrons consists of full s and p sub shells on an atom. In terms of Lewis symbols, an octet can be thought of as four pairs of valence electrons arranged around the atom as in the configuration for Ne, which is

There are many exceptions to the octet rule, but it provides a useful framework for many important bonding concepts.

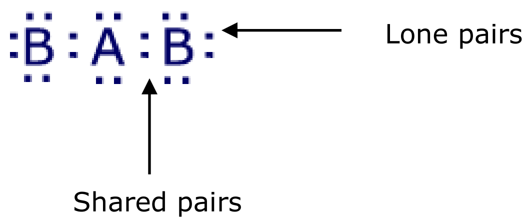
Skill 19.03: Explain the formation single, double, and triple covalent bonds

Skill 19.03 Concepts

A covalent bond is a bond in which two electrons are shared by two atoms. By sharing electrons in a covalent bond, the individual atoms can complete their octets and acquire greater stability.

- Covalent bonds result when two atoms similar electronegativity bond
- Covalent bonding between many-electron atoms involves only the valence electrons.

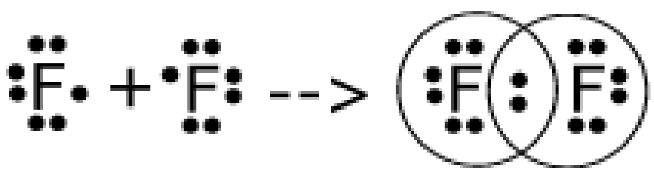
A Lewis structure is a representation of covalent bonding in which shared electron pairs are shown either as lines or as pairs of dots between two atoms, and lone pairs are shown as pairs of dots on individual atoms.



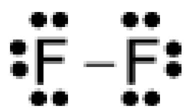
Atoms can form different types of covalent bonds:

- In a single bond, two atoms are held together by one electron pair.
- In a double bond, two atoms are held together by two electron pairs.
- In a triple bond, two atoms are held together by three electron pairs.

The formation of a bond between two F atoms to give a F_2 molecule can be represented in a similar way.

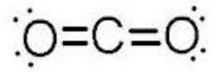


By sharing the bonding electron pair, each fluorine atom acquires eight electrons (an octet) in its valence shell. It thus achieves the noble-gas electron configuration of neon. The structures shown here for H_2 and F_2 are called Lewis structures (or Lewis electron-dot structures). In writing Lewis structures, we usually show each electron pair shared between atoms as a line, to emphasize a bond, and the unshared electron pairs as dots. Writing them this way, the Lewis structure for F_2 is shown as follows:



The number of valence electrons for the nonmetal is the same as the group number. Therefore, one might predict that 7A (group 17) elements such as F, would form one covalent bond to achieve an octet; 6A (group 16) elements, such as O, would form two covalent bonds; 5A (group 15), such as N, would form three covalent bonds; and 4A (group 14) elements, such as C, would form four covalent bonds.

The sharing of a pair of electrons constitutes a single covalent bond, generally referred to simply as a single bond. In many molecules, atoms attain an octet by sharing more than one pair of electrons between them. When two electron pairs are shared, two lines (representing a double bond) are drawn. A triple bond corresponds to the sharing of three pairs of electrons. Such multiple bonding is found in CO_2 and N_2 as shown below.

Skill 19.04: Draw Lewis structures for atoms that follow the octet rule

Skill 19.04 Concepts

A simple procedure for drawing Lewis structures is summarized below:

1. Count the total number of valence electrons present. For polyatomic anions, add the number of negative charges to that total. For polyatomic cations, subtract the number of positive charges from that total.

2. Write the skeletal structure of the compound, using the chemical symbols and placing bonded atoms next to one another. For compounds that contain more than two elements, typically the least electronegative atom occupies the central position.

3. Complete the octets of the atoms bonded to the central atom using the allowed electrons counted in step 1.

4. If the octet of the central atom is not complete after step 3, try adding double or triple bonds between the surrounding atoms and the central atom, using the lone pairs from the surrounding atoms. If there are left over electrons after the octets are complete, place them on the central atom.

Skill 19.04 Problem 1

Write the Lewis structures for the following molecules	
Cl_2	Br_2
O_2	N_2
CCl_4	SiF_4
OF_2	H_2O

CF_2H_2	CCl_2H_2
PF_3	NI_3
$COCl_2$	CO_2

The examples above all represented neutral molecules. Neutral molecules do not have a charge. Examples of molecules with a charge, also referred to as polyatomic ions, are shown below,

NO_3^-, CO_3^{2-}, PO_4^{3-}, etc

When drawing the Lewis Structures of molecules with a charge we must consider the charge in the total electron count. Below is example,

<u>Example</u>

Draw the Lewis structure for ClO_4^-

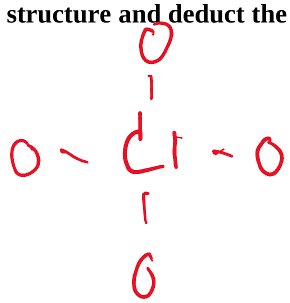
Step 1: Count the total valence electrons

Cl has 7 valence electrons, O has 6

The total valence electrons is therefore, 7 + 4(6) + 1 = 32

Notice above we multiplied oxygen by 4 because there are four oxygen atoms, we also added 1 to account for the 1- charge.

Step 2: Draw the skeleton structure and deduct the electrons used

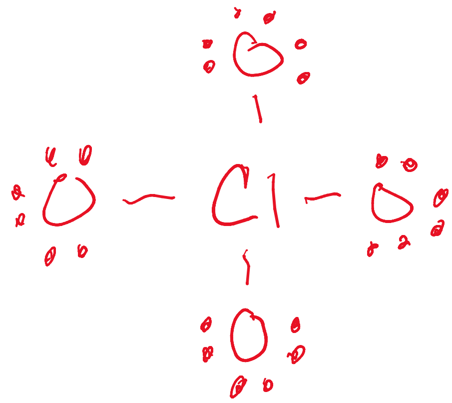


The skeleton structure above required 8 electrons – 2 for each bond. So, we will deduct this from allowed valance electrons from step 1,

32 – 8 = 24

This number, 24, is the number of electrons we are allowed to complete the structure.

Step 3: Complete the octets of the atoms bonded to the central atom



Before we fill the octets, we must remember that we cannot exceed the 24 electrons that we calculated in step 2

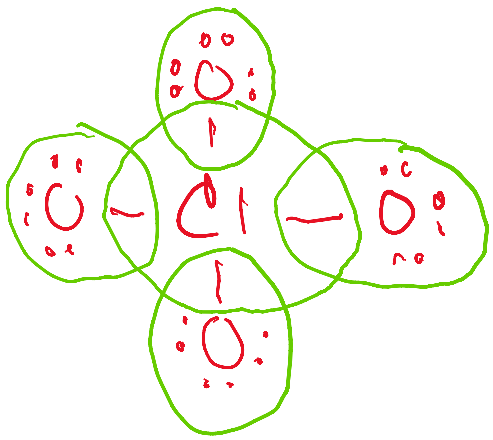
Each oxygen in the skeleton structure we drew in part 2 has 2 electrons. To complete the octets we have to add 6 more electrons to each oxygen. These electrons are represented as dots around the oxygens above. Notice we had to add 6 electrons to each oxygen for a total of 6(4) = 24.

Deducting the 24 electrons needed from the 24 that we were allowed gives us zero – we are out of electrons!

24 – 24 = 0

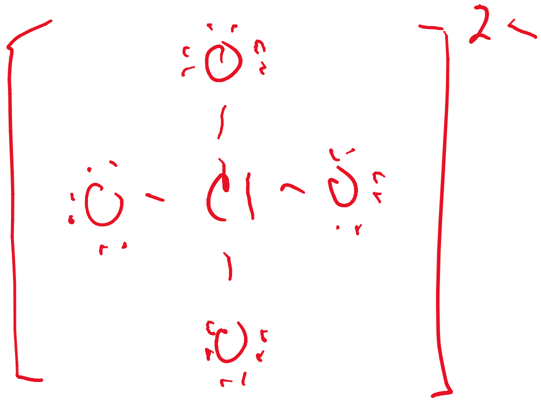
Step 5: Inspect the octets of all atoms in the structure

Now that we have used up our electrons, we must inspect the structure to see that each atom has the electrons they need.



The green circles represent the electrons around each atom. Chlorine has 8 because each bond represents 2 – 2(4) = 8. Each oxygen also has 8 – 1 bond + 6 unbonded electrons = 8.

Because ClO_4^{2-} is an ion, the final structure should be written as follows,



Skill 19.04 Problem 2

Write the Lewis structures for the following ions	
SO_4^{2-}	PO_4^{3-}
PO_4^{3-}	CN^-